12.83 An empty steel gas tank with valve weighs 125 lb. Its capacity is 1.50 ft^3. When the tank is filled with oxygen to 2000 lbf/in^2 at 25 °C, what percent of the total weight of the full tank is O_2? Assume the validity of the ideal gas law.

$$(125\ \text{lb})\left(\frac{1\ \text{kg}}{2.20\ \text{lb}}\right) = 56.8\ \text{kg} \qquad V = (1.50\ \text{ft}^3)\left(\frac{(12.0)(2.54)\ \text{cm}}{\text{ft}}\right)^3\left(\frac{1\ \text{L}}{1000\ \text{cm}^3}\right) = 42.5\ \text{L}$$

$$P = (2000\ \text{lbf/in.}^2)\left(\frac{1\ \text{atm}}{14.7\ \text{lbf/in.}^2}\right) = 136\ \text{atm} \qquad n = \frac{PV}{RT} = \frac{(136\ \text{atm})(42.5\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(298\ \text{K})} = 236\ \text{mol}$$

$$(236\ \text{mol})\left(\frac{0.032\ \text{kg}}{\text{mol}}\right)\left(\frac{2.2\ \text{lb}}{\text{kg}}\right) = 16.6\ \text{lb O}_2 \qquad \left(\frac{16.6\ \text{lb}}{(125\ \text{lb}) + (16.6\ \text{lb})}\right)(100\%) = 11.7\%$$

-

12.84 Pure oxygen gas is not necessarily the most compact source of the oxygen for confined fuel systems because of the weight of the cylinder necessary to confine the gas. Other compact sources are hydrogen peroxide and lithium peroxide. The oxygen-yielding reactions are:

$$2H_2O_2 \longrightarrow 2H_2O + O_2 \qquad 2Li_2O_2 \longrightarrow 2Li_2O + O_2$$

Rate (a) 65% (by weight) H_2O_2 and (b) pure Li_2O_2 in terms of % of total weight which is “available” oxygen. Neglect the weights of the containers.

(*a*) 1.00 kg sample contains 650 g H_2O_2.

$$(650\ \text{g H}_2\text{O}_2)\left(\frac{1\ \text{mol H}_2\text{O}_2}{34.0\ \text{g H}_2\text{O}_2}\right)\left(\frac{1\ \text{mol O}_2}{2\ \text{mol H}_2\text{O}_2}\right)\left(\frac{32.0\ \text{g O}_2}{\text{mol O}_2}\right) = 0.31 \times 10^3\ \text{g O}_2 = 0.31\ \text{kg O}_2$$

Compare with problem 12.83.

-

31% O_2 available

(b) $$(1.00 \times 10^3 \text{ g } Li_2O_2)\left(\frac{1 \text{ mol } Li_2O_2}{45.88 \text{ g } Li_2O_2}\right)\left(\frac{1 \text{ mol } O_2}{2 \text{ mol } Li_2O_2}\right)\left(\frac{32.0 \text{ g } O_2}{\text{mol } O_2}\right) = 0.35 \times 10^3 \text{ g } O_2 = 0.35 \text{ kg } O_2$$

35% O_2 available

12.85 What percent of a sample of nitrogen must be allowed to escape if its temperature, pressure, and volume are to be changed from 220 °C, 3.00 atm, and 1.65 L to 110 °C, 0.700 atm, and 1.00 L, respectively?

- $$n_1 = \frac{(3.00 \text{ atm})(1.65 \text{ L})}{R(493 \text{ K})} \qquad n_2 = \frac{(0.700 \text{ atm})(1.00 \text{ L})}{R(383 \text{ K})}$$

$$\text{Fraction remaining} = \frac{n_2}{n_1} = \frac{(0.700 \text{ atm})(1.00 \text{ L})/R(383 \text{ K})}{(3.00 \text{ atm})(1.65 \text{ L})/R(493 \text{ K})} = 0.182$$

$$\text{Fraction escaped} = 1.000 - 0.182 = 0.818 \qquad \text{Percent escaped} = 81.8\%$$

12.4 DALTON'S LAW

12.86 Show that the volume ratio of two gases, before mixing and at the same temperature and pressure, is equal to the pressure ratio after the two gases is mixed.

- Before mixing, T and P are the same for the two gases: $\frac{V_1}{V_2} = \frac{n_1RT/P}{n_2RT/P} = \frac{n_1}{n_2}$

After mixing, T and V are the same: $\frac{P_1}{P_2} = \frac{n_1RT/V}{n_2RT/V} = \frac{n_1}{n_2}$

From this proof, it is seen that the terms "volume percent" and "mole percent" are equivalent, and that the term "partial pressure percent" could also be used synonymously.

12.87 A mixture of gases at 760 torr contains 55.0% nitrogen, 25% oxygen, and 20.0% carbon dioxide by volume. What is the partial pressure of each gas in torr?

- A fundamental property of gases is that each component of a gas mixture occupies the entire volume of the mixture. The percent composition by volume refers to the separate gases *before* mixing. Thus 55.0 volumes of nitrogen, 25.0 volumes of oxygen, and 20.0

volumes of carbon dioxide, each at 760 torr, are mixed to give 100.0 volumes of mixture at 760 torr. As the volume of each component is increased by mixing, its pressure must (Boyle's law) decrease proportionately.

$$\text{Partial pressure of } N_2 = (760\text{ torr})(55.0/100) = 418\text{ torr}$$

$$\text{Partial pressure of } O_2 = (760\text{ torr})(25.0/100) = 190\text{ torr}$$

2 torr

0 torr

Check:

$$(1.00\text{ g } H_2)\left(\frac{1\text{ mol } H_2}{2.00\text{ g } H_2}\right) = 0.500\text{ mol } H_2 \qquad (88.0\text{ g } CO_2)\left(\frac{1\text{ mol } CO_2}{44.0\text{ g } CO_2}\right) = 2.00\text{ mol } CO_2$$

$$P(H_2) = \frac{n(H_2O)RT}{V} = \frac{(0.500\text{ mol})(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(300\text{ K})}{3.00\text{ L}} = 4.10\text{ atm}$$

12.88 A 1.00 L vessel containing 1.00 g H_2 gas at 27.0 °C is connected to a 2.00 L vessel containing 88.0 g CO_2 gas, also at 27.0 °C. When the gases are completely mixed, what are the partial pressures and total pressure (in atm)?

▪

$$P(CO_2) = \frac{(2.00\text{ mol})(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(300\text{ K})}{3.00\text{ L}} = 16.4\text{ atm} \qquad P(\text{total}) = (4.10\text{ atm}) + (16.4\text{ atm}) = 20.5\text{ atm}$$

12.89 In a gaseous mixture at 20 °C the partial pressures of the components are: hydrogen, 150 torr; carbon dioxide, 200 torr; methane, 320 torr; ethylene, 105 torr. What are the total pressure of the mixture and the volume percent of hydrogen?

$$\text{Total pressure of mixture} = \text{sum of partial pressures} = 150 + 200 + 320 + 105 = 775\text{ torr}$$

$$\text{Volume fraction of } H_2 = \frac{\text{partial pressure of } H_2}{\text{total pressure of mixture}} = \frac{150\text{ torr}}{775\text{ torr}} = 0.194 = 19.4\%$$

▪

12.90 A 200 mL flask contained oxygen at 220 torr, and a 300 mL flask contained nitrogen at 100 torr. The two flasks were then connected so that each gas filled their combined volume. Assuming no change in temperature, what was the partial pressure of each gas in the final mixture and what was the total pressure?

- The final total volume was 500 mL.

Oxygen:
$$P_{\text{final}} = P_{\text{initial}}\left(\frac{V_{\text{initial}}}{V_{\text{final}}}\right) = (220\text{ torr})\left(\frac{200}{500}\right) = 88\text{ torr}$$

Nitrogen:
$$P_f = P_i\left(\frac{V_i}{V_f}\right) = (100\text{ torr})\left(\frac{300}{500}\right) = 60\text{ torr}$$

$$\text{Total pressure} = (88\text{ torr}) + (60\text{ torr}) = 148\text{ torr}$$

12.91 If 3.00 L oxygen gas is collected over water at 27°C when the barometric pressure is 787 torr, what is the volume of dry gas at 0°C and 1.00 atm, assuming ideal behavior? Vapor pressure of water at 27°C is 27 torr.

$$P(O_2) = (787\text{ torr}) - (27\text{ torr}) = 760\text{ torr} = 1.00\text{ atm} \qquad n(O_2) = \frac{PV}{RT} = \frac{(1.00)(3.00)}{(0.0821)(300)} = 0.122\text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(0.122\text{ mol})(0.0821\text{ L·atm/mol·K})(273\text{ K})}{1.00\text{ atm}} = 2.73\text{ L}$$

-

12.92 Exactly 100 mL of oxygen is collected over water at 23 °C and 800 torr. Compute the standard volume of the dry oxygen. Vapor pressure of water at 23 °C is 21.1 torr.

- The gas collected is actually a mixture of oxygen and water vapor. The partial pressure of water vapor in the mixture at 23 °C is 21.1 torr. Hence

Pressure of dry oxygen = (total pressure) − (vapor pressure of water) = (800 torr) − (21 torr) = 779 torr

Thus, for the dry oxygen, $V_1 = 100\ \text{mL}$, $T_1 = 23 + 273 = 296\ \text{K}$, $P_1 = 779\ \text{torr}$. Converting to STP,

$$V_2 = V_1\left(\frac{T_2}{T_1}\right)\left(\frac{P_1}{P_2}\right) = (100\ \text{mL})\left(\frac{273\ \text{K}}{296\ \text{K}}\right)\left(\frac{779\ \text{torr}}{760\ \text{torr}}\right) = 94.5\ \text{mL}$$

12.93 In a basal metabolism measurement timed at 6.00 min, a patient exhaled 52.5L of air, measure over water at 20 °C. The vapor pressure of water at 20 °C is 17.5 torr. The barometric pressure was 750 torr. The exhaled air analyzed 16.75 vol% oxygen, and the inhaled air 20.32 vol% oxygen, both on a dry basis. Neglecting any solubility of the gases in water and any difference in the total volumes of inhaled and exhaled air, calculate the rate of oxygen consumption by the patient in mL (STP) per minute.

- $$\text{Volume of dry air at STP} = (52.5\ \text{L})\left(\frac{273\ \text{K}}{293\ \text{K}}\right)\left[\frac{(750 - 17.5)\ \text{torr}}{760\ \text{torr}}\right] = 47.1\ \text{L}$$

$$\text{Rate of oxygen consumption} = \frac{\text{volume of oxygen (STP) consumed}}{\text{time in which this volume was consumed}}$$

$$= \frac{(0.2032 - 0.1675)(47.1\ \text{L})}{6.00\ \text{min}} = 0.280\ \text{L/min} = 280\ \text{mL/min}$$

12.94 Exactly 0.550 L nitrogen is collected over water at 25 °C and 755 torr. The gas is saturated with water vapor. Compute the volume of the nitrogen in the dry condition at STP. Vapor pressure of water at 25 °C is 23.8 torr.

$$P(N_2) = P(\text{total}) - P(H_2O) = (755\ \text{torr}) - (23.8\ \text{torr}) = 731\ \text{torr}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2} = \frac{(0.550\ \text{L})(731\ \text{torr})(273\ \text{K})}{(298\ \text{K})(760\ \text{torr})} = 0.485\ \text{L}$$

12.95 A dry gas occupied 150 mL at STP. If this same mass of gas were collected over water at 23 °C and a total gas pressure of 745 torr, what volume would it occupy? The vapor pressure of water at 25 °C is 23.8 torr.

- The final pressure of the dry gas must be used. (imagine expanding the dry gas to V_2 and then adding the water vapor, which increases the pressure but not the volume.)

$$P(\text{gas}) = P(\text{total}) - P(H_2O) = (745\ \text{torr}) - (21\ \text{torr}) = 724\ \text{torr}$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad V_2 = \frac{P_1V_1T_2}{T_1P_2} = \frac{(760\ \text{torr})(150\ \text{mL})(296\ \text{K})}{(273\ \text{K})(724\ \text{torr})} = 171\ \text{mL}$$

12.96 A mixture of N_2, NO, and NO_2 was analyzed by selective absorption of the oxides of the nitrogen. The initial volume of the mixture was 2.74 cm^3. After treatment with water, which absorbed the NO_2, the volume was 2.02 cm^3. A ferrous sulfate solution was then shaken with the residual gas to absorb the NO, after which the volume was 0.25 cm^3. All volumes were measured at barometric pressure. Neglecting water vapor, what was the volume percent of each gas in the original mixture?

- $$\%\ N_2 = \left(\frac{0.25\ \text{cm}^3}{2.74\ \text{cm}^3}\right)(100\%) = 9.1\%\ N_2 \qquad \%\ NO = \left(\frac{(2.02 - 0.25)\ \text{cm}^3}{2.74\ \text{cm}^3}\right)(100\%) = 64.6\%\ NO$$

$$\%\ NO_2 = \left(\frac{(2.74 - 2.02)\ \text{cm}^3}{2.74\ \text{cm}^3}\right)(100\%) = 26.3\%\ NO_2$$

12.97 A 250 mL flask contained krypton at 500 torr. A 450 mL flask contained helium at 950 torr. The contents of the two flasks were mixed by opening a stopcock connecting them. Assuming that all operations were carried out at a uniform constant temperature, calculate the final total pressure and the volume percent of each gas in the mixture. Neglect the volume of the stopcock.

- $$P_2(\text{Kr}) = P_1(\text{Kr})\left(\frac{V_1}{V_2}\right) = (500\text{ torr})\left(\frac{250\text{ mL}}{700\text{ mL}}\right) = 179\text{ torr} \qquad P_2(\text{He})\left(\frac{V_1}{V_2}\right) = (950\text{ torr})\left(\frac{450\text{ mL}}{700\text{ mL}}\right) = 611\text{ torr}$$

  $$\text{Total pressure} = 790\text{ torr} \qquad \text{Vol }\% = \left(\frac{179\text{ torr}}{790\text{ torr}}\right)(100\%) = 22.7\%\text{ Kr} \qquad \text{and} \qquad 77.3\%\text{ He}$$

12.98 The vapor pressure of water at 80 °C is 355 torr. A 100 mL vessel contained water-saturated oxygen at 80 °C, the total gas pressure being 760 torr. The contents of the vessel were pumped into a 50.0 mL vessel at the same temperature. What were the partial pressures of oxygen and of water vapor, and what was the total pressure in the final equilibrated state? Neglect the volume of any water which might condense.

- The partial pressure of water remains the same, 355 torr; the excess water condenses. The partial pressure of oxygen doubles, 2(760 -355) = 810 torr. The total pressure is just the sum, 1165 torr.

12.99 Show that for mixture of gases P(total)V = n(total)RT.

- The total number of moles of gas is given by

  $$n(\text{total}) = n_1 + n_2 + \cdots = \frac{P_1V_1}{RT_1} + \frac{P_2V_2}{RT_2} + \cdots$$

Since all the gases in the mixture are in the same container, they all occupy the same volume and they all have the same temperature:

$$n(\text{total}) = \frac{P_1V}{RT} + \frac{P_2V}{RT} + \cdots = \frac{V}{RT}(P_1 + P_2 + \cdots)$$

Thus $n(\text{total})RT = V(P_1 + P_2 + \cdots) = VP(\text{total})$

12.100 Oxygen is collected over water at 25 °C in a 2.00 L vessel at a total barometric pressure of 765 torr. Calculate the number of moles of oxygen collected.

- The vapor pressure of water at 25 °C is 24 torr. Thus the pressure due to the oxygen is (765 torr) – (24 torr) = 741 torr. The number of moles of oxygen maybe calculated:

$$n = \frac{PV}{RT} = \frac{(\frac{741}{760}\ \text{atm})(2.00\ \text{L})}{(0.0821\ \text{L·atm/mol·K})(298\ \text{K})} = 0.0797\ \text{mol}$$

12.101 If 2.0 g N_2, 0.40 g H_2, and 9.0 g O_2 are put into a 1.00 L container at 27 °C, what is the total pressure in the container?

- See Problem 12.99.

$$n(\text{total}) = \left(\frac{2.0\ \text{g}}{28.0\ \text{g/mol}}\right) + \left(\frac{0.40\ \text{g}}{2.0\ \text{g/mol}}\right) + \left(\frac{9.0\ \text{g}}{32.0\ \text{g/mol}}\right) = 0.55\ \text{mol}$$

$$P(\text{total}) = \frac{(0.55\ \text{mol})(0.0821\ \text{L·atm/mol·K})(300\ \text{K})}{1.00\ \text{L}} = 14\ \text{atm}$$

12.102 A 15.0 L vessel containing 5.65 g N_2 is connected by means of valve to a 6.00 L vessel containing 5.00 g oxygen. After the valve is opened and the gases are allowed to mix, what will be the partial pressure of each gas and the total pressure at 27 °C?

- The total volume after mixing is 21.0 L

$$(5.65\ \text{g}\ N_2)\left(\frac{1\ \text{mol}\ N_2}{28.0\ \text{g}}\right) = 0.202\ \text{mol}\ N_2 \qquad (5.00\ \text{g}\ O_2)\left(\frac{1\ \text{mol}\ O_2}{32.0\ \text{g}}\right) = 0.156\ \text{mol}\ O_2$$

$$P(N_2) = \frac{nRT}{V} = \frac{(0.202\ \text{mol})(0.0821\ \text{L·atm/mol·K})(300\ \text{K})}{21.0\ \text{L}} = 0.237\ \text{atm}$$

$$P(O_2) = \frac{(0.156\ \text{mol})(0.0821\ \text{L·atm/mol·K})(300\ \text{K})}{21.0\ \text{L}} = 0.183\ \text{atm}$$

$$P(\text{total}) = P(N_2) + P(O_2) = (0.237\ \text{atm}) + (0.183\ \text{atm}) = 0.420\ \text{atm}$$

12.103 Into a 5.00 L container at 18 °C are placed 0.200 mol H_2,

$$P(H_2) = \frac{nRT}{V} = \frac{(0.200\ \text{mol})(0.0821\ \text{L·atm/mol·K})(291\ \text{K})}{5.00\ \text{L}} = 0.956\ \text{atm}$$

$$n(CO_2) = (20.0\ \text{g})\left(\frac{1\ \text{mol}}{44.0\ \text{g}}\right) = 0.4545\ \text{mol}\ CO_2$$

$$P(CO_2) = \frac{(0.4545\ \text{mol})(0.0821\ \text{L·atm/mol·K})(291\ \text{K})}{5.00\ \text{L}} = 2.17\ \text{atm}$$

$$n(O_2) = (14.00\ \text{g})\left(\frac{1\ \text{mol}}{32.00\ \text{g}}\right) = 0.4375\ \text{mol}\ O_2$$

$$P(O_2) = \frac{(0.4375\ \text{mol})(0.0821\ \text{L·atm/mol·K})(291\ \text{K})}{5.00\ \text{L}} = 2.09\ \text{atm}$$

$$P(\text{total}) = (0.956\ \text{atm}) + (2.17\ \text{atm}) + (2.09\ \text{atm}) = 5.22\ \text{atm}$$

20.0 g CO_2, and 14.00 g O_2. Calculate the total pressure in the container and the partial pressure of each gas.

▪

12.104 How many g off oxygen is contained in 10.5 L of oxygen measures over water at 25 °C and 740 torr? Vapor pressure of water at 25 °C is 24 torr.

$$P(O_2) = P(\text{total}) - P(\text{water}) = (740\ \text{torr}) - (24\ \text{torr}) = 716\ \text{torr}$$

$$n = \frac{PV}{RT} = \frac{(\frac{716}{760}\ \text{atm})(10.5\ \text{L})}{(0.0821\ \text{L·atm/mol·K})(298\ \text{K})} = 0.404\ \text{mol} \qquad (0.404\ \text{mol}\left(\frac{32.0\ \text{g}}{\text{mol}}\right) = 12.9\ \text{g}$$

▪

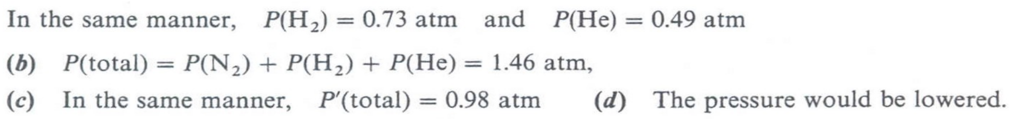
12.105 If 0.10 mol N_2, 0.300 mol H_2, and 0.20 mol He are placed in a 10.0 L container at 25 °C, (a) what are the volume, temperature, and pressure of each gas? (b) What is the total pressure? (c) if 0.20 mol NH_3 and 0.20 mol of He are placed in a similar container at the same temperature, what is the the total pressure in container? (d) Would a reaction

(a) $V = 10.0\ \text{L}$ for each; $T = 298\ \text{K}$ for each.

$$P(N_2) = \frac{nRT}{V} = \frac{(0.10\ \text{mol})(0.0821\ \text{L·atm/mol·K})(298\ \text{K})}{10.0\ \text{L}} = 0.24\ \text{atm}$$

converting some N_2 and H_2 into NH_3 raise or lower the total pressure in the first container?

-

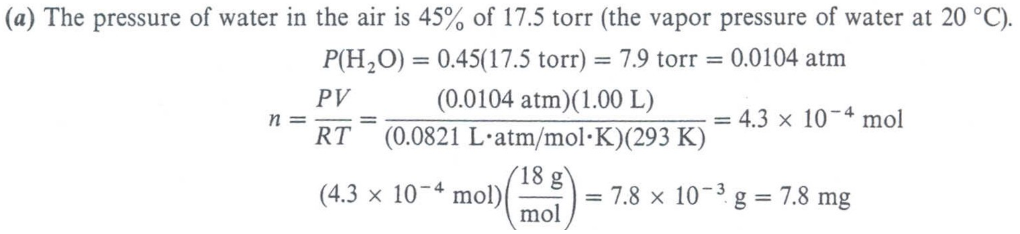


In the same manner, $P(H_2) = 0.73$ atm and $P(He) = 0.49$ atm

(*b*) $P(\text{total}) = P(N_2) + P(H_2) + P(He) = 1.46$ atm,

(*c*) In the same manner, $P'(\text{total}) = 0.98$ atm (*d*) The pressure would be lowered.

12.106 Relative humidity is defined as the ratio of the partial pressure of water in air at a given temperature to the vapor pressure of water at that temperature. Calculate the mass of water per liter of air at (a) 20 °C and 45% relative humidity, (b) 0 °C and 95% relative humidity. (c) Discuss whether temperature or relative humidity has the greater effect on the mass of water vapor in the air.



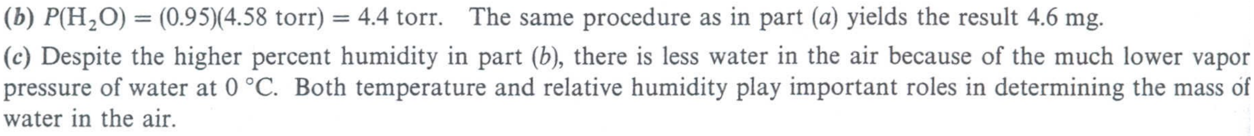
(*a*) The pressure of water in the air is 45% of 17.5 torr (the vapor pressure of water at 20 °C).

$$P(H_2O) = 0.45(17.5 \text{ torr}) = 7.9 \text{ torr} = 0.0104 \text{ atm}$$

$$n = \frac{PV}{RT} = \frac{(0.0104 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L·atm/mol·K})(293 \text{ K})} = 4.3 \times 10^{-4} \text{ mol}$$

$$(4.3 \times 10^{-4} \text{ mol})\left(\frac{18 \text{ g}}{\text{mol}}\right) = 7.8 \times 10^{-3} \text{ g} = 7.8 \text{ mg}$$

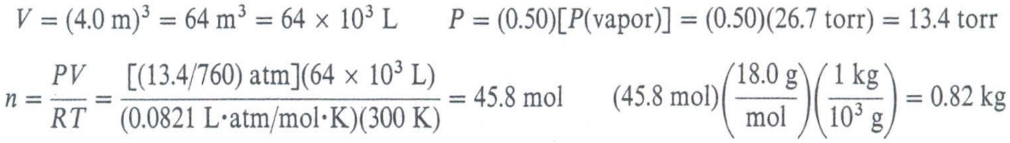
-



(*b*) $P(H_2O) = (0.95)(4.58 \text{ torr}) = 4.4$ torr. The same procedure as in part (*a*) yields the result 4.6 mg.

(*c*) Despite the higher percent humidity in part (*b*), there is less water in the air because of the much lower vapor pressure of water at 0 °C. Both temperature and relative humidity play important roles in determining the mass of water in the air.

12.107 How much water vapor is contained in a cubic room 4.0 m along an edge if the relative humidity is 50% and the temperature is 27 °C? The vapor



$$V = (4.0 \text{ m})^3 = 64 \text{ m}^3 = 64 \times 10^3 \text{ L} \qquad P = (0.50)[P(\text{vapor})] = (0.50)(26.7 \text{ torr}) = 13.4 \text{ torr}$$

$$n = \frac{PV}{RT} = \frac{[(13.4/760) \text{ atm}](64 \times 10^3 \text{ L})}{(0.0821 \text{ L·atm/mol·K})(300 \text{ K})} = 45.8 \text{ mol} \qquad (45.8 \text{ mol})\left(\frac{18.0 \text{ g}}{\text{mol}}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right) = 0.82 \text{ kg}$$

pressure of water at 27 °C is 26.7 torr. (the relative humidity expresses the partial pressure of water as a percent of the water vapor pressure.)

-

12.5 MOLECULAR WEIGHTS OF GASES

12.108 Avogadro’s law states that equal volumes of gases under the same conditions of temperature and pressure contain equal numbers of molecules. Develop from that law the fact that at a given temperature and pressure, the density of a pure gas is proportional to its molecular weight.

- $PV = nRT = \frac{m}{\mathrm{MW}}RT$ where m is the mass of the gas and MW is the molecular weight.

Rearranging algebraically yields $$\frac{m}{V} = \frac{P}{RT}(\mathrm{MW})$$

Since m/V is the density of the gas, and under the conditions stated in the problem T and P are constant, the equation can be restated Density = const. x MW which is a mathematical statement of direct proportionality.

12.109 Determine the approximate molecular weight of a gas if 560 mL weighs 1.80 g at STP.

$$PV = \left(\frac{m}{\mathrm{MW}}\right)RT \qquad \mathrm{MW} = \frac{mRT}{PV} = \frac{(1.80\ \mathrm{g})(0.0821\ \mathrm{L\cdot atm/K\cdot mol})(273\ \mathrm{K})}{(1.00\ \mathrm{atm})(0.560\ \mathrm{L})} = 72.0\ \mathrm{g/mol}$$

-

12.110 At 18 °C and 765 torr, 2.29 L of a gas weighs 3.71 g. calculate the approximate molecular weight of the gas.

Converting the data to atm and K, $P = \frac{765}{760}\ \mathrm{atm}$ $T = (18 + 273)\ \mathrm{K} = 291\ \mathrm{K}$

-

Then $$\mathrm{MW} = \frac{mRT}{PV} = \frac{(3.71\ \mathrm{g})(0.0821\ \mathrm{L\cdot atm/K\cdot mol})(291\ \mathrm{K})}{(\frac{765}{760}\ \mathrm{atm})(2.29\ \mathrm{L})} = 38.5\ \mathrm{g/mol}$$

12.111 If the density of carbon monoxide is 3.17 g/L at -20 °C and 2.35 atm, what is its approximate molecular weight?

$$\mathrm{MW} = \frac{dRT}{P} = \frac{(3.17\ \mathrm{g/L})(0.0821\ \mathrm{L\cdot atm/K\cdot mol})(253\ \mathrm{K})}{2.35\ \mathrm{atm}} = 28.0\ \mathrm{g/mol}$$

-

12.112 At 750 torr and 27 °C, 0.60 g of certain gas occupies 0.50 L. Calculate its molecular weight.

$$PV = nRT \qquad (\tfrac{750}{760}\ \mathrm{atm})(0.50\ \mathrm{L}) = n\left(0.0821\ \frac{\mathrm{L\cdot atm}}{\mathrm{mol\cdot K}}\right)(300\ \mathrm{K}) \qquad n = 2.0 \times 10^{-2}\ \mathrm{mol}$$

- $$\mathrm{MW} = \frac{\text{mass}}{\text{number of moles}} = \frac{0.60\ \mathrm{g}}{2.0 \times 10^{-2}\ \mathrm{mol}} = 30\ \mathrm{g/mol}$$

12.113 Calculate the MW of a gas which has a density of 1.48 g/L at 100 °C *and* 600 torr.

Per liter:

$$n = \frac{PV}{RT} = \frac{(\frac{600}{760}\ \mathrm{atm})(1.00\ \mathrm{L})}{(0.0821\ \mathrm{L\cdot atm/mol\cdot K})(373\ \mathrm{K})} = 0.0258\ \mathrm{mol}$$

- $$\mathrm{MW} = \frac{1.48\ \mathrm{g/L}}{0.0258\ \mathrm{mol/L}} = 57.4\ \mathrm{g/mol}$$

12.114 If 300 mL of gas weighs 0.368 g at STP, what is its molecular weight?

- $$n = \frac{PV}{RT} = \frac{(1.00\ \mathrm{atm})(0.300\ \mathrm{L})}{(0.0821\ \mathrm{L\cdot atm/mol\cdot K})(273\ \mathrm{K})} = 13.4\ \mathrm{mmol}$$

$$\mathrm{MW} = \frac{368\ \mathrm{mg}}{13.4\ \mathrm{mmol}} = 27.5\ \mathrm{mg/mmol} = 27.5\ \mathrm{g/mol}$$

12.115 At 27 °C, 2.40 g of certain gas occupies a volume of 9.84 L at a pressure of 172 torr. Calculate the MW of the gas.

$$n = \frac{PV}{RT} = \frac{[(172/760)\text{ atm}](9.84\text{ L})}{(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(300\text{ K})} = 0.0904\text{ mol} \qquad \frac{2.40\text{ g}}{0.0904\text{ mol}} = 26.5\text{ g/mol}$$

▪

12.116 A 2.00 g sample of gas at 27 °C and 1.00 atm occupies 1.00 L. What is the MW of the gas?

$$n = \frac{PV}{RT} = \frac{(1.00\text{ atm})(1.00\text{ L})}{(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(300\text{ K})} = 0.0406\text{ mol} \qquad \text{MW} = \frac{2.00\text{ g}}{0.0406\text{ mol}} = 49.3\text{ g/mol}$$

▪

12.117 A 1.525 g mass of a volatile liquid is vaporized, giving 500 cm^3 of vapor when measured over water at 30 °C and 770 torr. Thee vapor pressure of water at 30 °C is 32 torr. What is the MW of the substances>

For the vapor,

$$P = (770\text{ torr}) - (32\text{ torr}) = (738\text{ torr})\left(\frac{1\text{ atm}}{760\text{ torr}}\right) = 0.971\text{ atm}$$

$$n = \frac{PV}{RT} = \frac{(0.971\text{ atm})(0.500\text{ L})}{(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(303\text{ K})} = 0.0195\text{ mol} \qquad \text{MW} = \frac{1.525\text{ g}}{0.0195\text{ mol}} = 78.2\text{ g/mol}$$

▪

12.118 Argon gas liberated from crushed meteorites does not have the same isotopic composition as atmosphere argon. The gas density of a particular sample of meteoritic Ar was found to be 1.481 g/L at 27 °C and 740 torr. What is the average atomic weight of this sample of argon?

- One liter of sample contains

$$n = \frac{PV}{RT} = \frac{\left[\frac{740}{760}\ \text{atm}\right](1.00\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(300\ \text{K})} = 0.0395\ \text{mol} \qquad \text{AW} = \frac{1.481\ \text{g/L}}{0.0395\ \text{mol/L}} = 37.5\ \text{g/mol}$$

12.119 The density of an unknown gas at 0.821 atm and 27 °C is 2.50 g/L.

$$\text{In 1.00 L:} \qquad n = \frac{PV}{RT} = \frac{(0.821\ \text{atm})(1.00\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(300\ \text{K})} = 0.0333\ \text{mol} \qquad \text{MW} = \frac{2.50\ \text{g/L}}{0.0333\ \text{mol/L}} = 75.0\ \text{g/mol}$$

What is the MW of this gas? (Assume ideal behavior.)

-

12.120 Find the MW of a gas if 75.0 g of this gas exerts a pressure of 2.00 atm in a 30.0 L vessel at 500 K.

$$n = \frac{PV}{RT} = \frac{(2.00\ \text{atm})(30.0\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(500\ \text{K})} = 1.46\ \text{mol} \qquad \text{MW} = \frac{75.0\ \text{g}}{1.46\ \text{mol}} = 51.3\ \text{g/mol}$$

-

12.121 The density of NO was very carefully determined to be 0.2579 kg/m^3 at a temperature and pressure at which oxygen's measured density was 0.2749 kg/m^3. On the basis of this information and the known atomic weight of oxygen, calculate the atomic weight of nitrogen.

- Density and molecular weight are proportional (Problem 12.108);

$$\frac{d_1}{\text{MW}_1} = \frac{d_2}{\text{MW}_2} \qquad \frac{0.2579\ \text{g/L}}{\text{MW}_1} = \frac{0.2749\ \text{g/L}}{32.00\ \text{g/mol}} \qquad \text{MW}_1 = (32.00\ \text{g/mol})\left(\frac{0.2579}{0.2749}\right) = 30.02\ \text{g/mol NO}$$

$$\text{AW} = (30.02\ \text{g/mol}) - (16.00\ \text{g/mol}) = 14.02\ \text{g/mol} = 14.02\ \text{u}$$

hence

12.122 (a) Assuming that air is 79 mol% N_2, 20% O_2, and 1% Ar, compute the average MW of air. (b) What is the density of air at 25 °C and 1 atm?

- $$(a)\quad \mathrm{MW} = (0.79)(28.0) + (0.20)(32.0) + (0.01)(39.9) = 28.9\ \mathrm{g/mol}$$
  $$(b)\quad d = \frac{(\mathrm{MW})P}{RT} = \frac{(28.9\ \mathrm{g/mol})(1.00\ \mathrm{atm})}{(0.0821\ \mathrm{L\cdot atm/mol\cdot K})(298\ \mathrm{K})} = 1.18\ \mathrm{g/L}$$

12.123 Devise a method fro determining the molecular weight of a low-boiling liquid.

- Heat the sample to vaporize the liquid at a know temperature and pressure in a vessel of known mass and volume, displacing all other gases and making sure that all liquid phase has evaporated. Cool the vessel, allowing the gas to condense to liquid. Weigh the vessel and its contents. The data now available are P,V,T and m. From the first three of these, the number of moles can be calculated. Then the MW is equal to m/n.

12.124 Find the MW of a gas whose density at 40 °C and 758 torr is 1.3866 kg/m³.

$$\mathrm{MW} = \frac{mRT}{PV} = \left(\frac{m}{V}\right)\left(\frac{RT}{P}\right) = d\left(\frac{RT}{P}\right)$$

- $$\left(\frac{1.386\ \mathrm{kg}}{\mathrm{m}^3}\right)\left(\frac{1000\ \mathrm{g}}{\mathrm{kg}}\right)\left(\frac{1\ \mathrm{m}}{10\ \mathrm{dm}}\right)^3 = 1.386\ \mathrm{g/dm^3} = 1.386\ \mathrm{g/L}$$
  $$\mathrm{MW} = \frac{(1.386\ \mathrm{g/L})(0.0821\ \mathrm{L\cdot atm/mol\cdot K})(313\ \mathrm{K})}{(758/760)\ \mathrm{atm}} = 35.7\ \mathrm{g/mol}$$

12.125 A gaseous compound has an empirical formula CH_3. Its density is 2.06 g/L at the same temperature and pressure at which oxygen has a density of 2.21 g/L. calculate the molecular formula of the gas.

- Rearrange PV =nRT to the form n/V = P/RT. Since P and T are the same for the oxygen and the unknown gas (x), their number of mol/L are also equal. Per liter:

$$n(O_2) = n(X) \qquad \frac{2.21\ \text{g}}{32.0\ \text{g/mol}} = \frac{2.06\ \text{g}}{\text{MW(X)}}$$

$$\text{MW(X)} = \left(\frac{2.06}{2.21}\right)\left(\frac{32.0\ \text{g}}{\text{mol}}\right) = 29.8\ \text{g/mol}$$

The empirical formula, CH_3, has a formula weight of 15 g/mol; so there must be 29.8/15 ≈ 2 formula units in a molecule of X. the molecular formula is $(CH_3)_2$ or C_2H_6.

12.126 A gas is composed of 30.4% N 69.6% O. its density is 11.1 g/L at -20 °C and 2.50 atm. What are the empirical and molecular formulas of the gas?

$$(30.4\ \text{g N})\left(\frac{1\ \text{mol N}}{14.0\ \text{g N}}\right) = 2.17\ \text{mol N} \qquad (69.6\ \text{g O})\left(\frac{1\ \text{mol O}}{16.0\ \text{g O}}\right) = 4.35\ \text{mol O}$$